Corrosion

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It is an electrochemical process where metals spontaneously oxidize to produce new substances by combining with oxygen or other non metals. Examples are rusting of iron, tarnishing of silver, formation of copper patina etc.

Corrosion happens when metal atoms lose electrons to become positive ions or get oxidized.

Rusting of Iron

There are two half reactions for the full reaction of rust to occur.

Note that both Oxygen and water are required for the rusting of Fe

$$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$$

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-$$

1. Fe turns into Fe^{2+} at a wet exposed surface
2. The electrons mix with the oxygen dissolved in the water
3. Makes Iron(II) Hydroxide which is further oxidized to form Iron (III) Hydroxide
4. Iron (III) Hydroxide dehydrates to Iron III oxide (rust)

Corrosion Prevention

Barrier Protection

- Painting the exterior
- Electroplating or placing other metals on the exterior of the metal being protected
- Greasing / Oiling

Cathodic Protection: Cause the Fe to be the cathode

- Galvanization: Using a more electropositive metal like zinc to be the anode and the metal to be protected usually Fe be the cathode. If the combination is Zn and Fe we say the iron is galvanized.
- Sacrificial protection: The metal to be protected is made the cathode and a more reactive metal is connected by an external wire usually magnesium is the anode. Example: Railway tracks, Ships, oil pipes connected to a chunk of magnesium using an external wire

Electroplating

Electroplating is a method used to cover the surface of an object with a thin layer of metal. This is a process that can take place by placing a piece of metal to be coated on the cathode of a voltaic cell. Can be explained using Faradays laws

Faraday's Law

- The relationship between electricity and electrochemical changes was first investigated by Michael Faraday in the 1830s.
- He discovered that the mass of an element produced or consumed at an electrode was directly proportional to the time the cell operated, as long as the current was constant.
- The charge of every mole of electrons that flows in the cell= 9.65 x 10^4 C/mol

$$n_e = \frac{It}{F}$$

$$Q = It$$

$$n_e = number\ of\ moles\ of\ electron$$

$$I = current\ strength\ in\ amperes$$

$$F\ or\ Faradays\ Constant = \ 6.022x10^{23}x1.6x10^{-19}C$$

$$F = 96500Coulumbs/mole$$

Problems

1. What amount of electrons is transferred in a cell that operates for 1.25 h at a current of 0.150 A?
2. Convert a current of 1.74 A for 10.0 min into an amount in moles of electrons?
3. How long, in minutes, will it take a current of 3.5o A to transfer 0.100 mol of electrons?
4. What is the mass of copper deposited at the cathode of a copper electro refining cell operated at 12.0 A for 40.0 min?
5. A student reconstructs electric battery using sheets of copper and zinc, and a current of 0.500 A is produced for 10.0 min. Calculate the mass of zinc oxidized to aqueous zinc ions.
6. Electroplating is a common technological process for coating objects with a metal to enhance the appearance of the object or its resistance to corrosion. A car bumper is plated with chromium using chromium (III) ions in solution. If a current of 54 A flows in the cell for 45 min 30 s, determine the mass of chromium deposited on the bumper.

Balance all half cells and the net ionic equation before calculating the quantities.